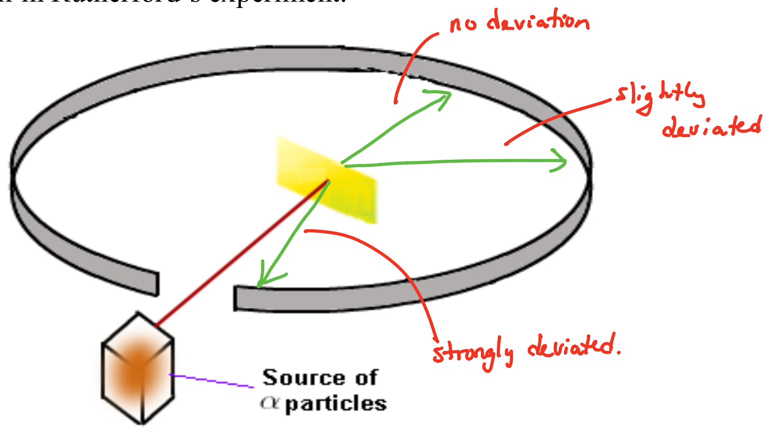
Name:_______________________________________ Date:______________________

SCH4U - Structures and Properties
Chapter 3 Test

Question 1.

a) Using arrows, show on the diagram below all of the possible pathways taken by the alpha particles as they went through the gold foil in Rutherford's experiment.



b) Explain how the different pathways in the diagram allowed Rutherford to develop his atomic model.

① α-particles that did not deviate or slightly deviated
↳ helped conclude that atom is+ mostly empty space.

② α-particles that strongly deviated
↳ helped conclude that there is a very small positive region at the centre of the atom (nucleus)

Question 2. How did Bohr's atomic model explain hydrogen's line spectrum?

① Electrons exist in a series of "allowable" energy levels.

② Electrons can move from one allowable energy level to another.

③ The energy absorbed and released corresponds to the energy difference between allowable energy levels

↳ Only particular wavelengths will be observed (as is the case for hydrogen)

Question 3. What information about the electron does the quantum mechanical model of the atom provide?

① Electrons as a standing wave

② Orbitals → region in which an electron spends 90% of its time.

③ Heisenberg uncertainty principle: impossible to know simultaneously with exact precision the position and momentum of an electron.

④ Quantum numbers ("address" of e^-)
n → energy level
l → orbital shape
m_l → oriention of orbital
m_s → e^- spin.

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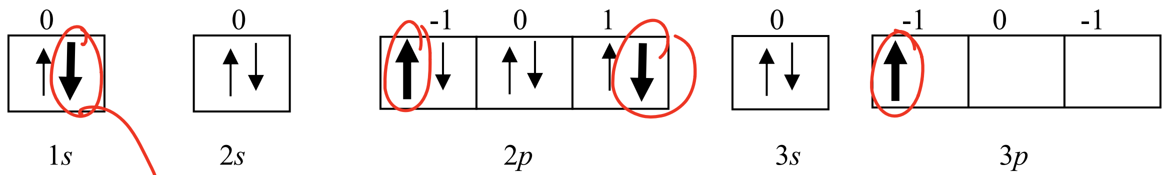
Question 4. Below is the full orbital diagram of an element at its ground-state. Provide the quantum numbers for the electrons in **bold** and identify the element.



(The electrons are numbered from left to right).

- Electron 1: $n=1$, $l=0$, $m_l=0$, $m_s=-1/2$
- Electron 2: $n=2$, $l=1$, $m_l=-1$, $m_s=+1/2$
- Electron 3: $n=2$, $l=1$, $m_l=1$, $m_s=-1/2$
- Electron 4: $n=3$, $l=1$, $m_l=-1$, $m_s=+1/2$
- Element: Aluminimum.

Question 5. Identify whether the following combinations of quantum numbers are permissible. If they are not, explain why.

a. $n = 2$, $l = 0$, $m_l = -1$, $m_s = +1/2$ not permissible, m_l can only be 0

b. $n = 3$, $l = 1$, $m_l = 0$, $m_s = +1/2$ permissible

c. $n = 10$, $l = 5$, $m_l = -3$, $m_s = -1/2$ permissible

d. $n = 2$, $l = 3$, $m_l = 0$, $m_s = +1/2$ not permissible l goes from 0 to $n-1$

Question 6.

a) Write the full electron configuration of As, N, and P.

As: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^3$

N: $1s^2 2s^2 2p^3$

P: $1s^2 2s^2 2p^6 3s^2 3p^3$

b) Predict the ranking of the atomic radius of As, N, and P. Justify your answer.

smallest to largest

N P As

① Trend: down a group atomic radius ↑

② increase in "n" starting from N.

③ shielding effect.

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Question 7.

a) Write the condensed electronic configuration of Germanium.

$[Ar]\ 4s^2 3d^{10} 4p^2$

b) Using Germanium as an example, explain how the electronic configuration can be used to identify the group number, the period number, the orbital block, and the total number of electrons of an element.

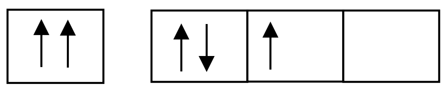
① group # : if in s-bock simply look at # of electrons in the highest s orbital.
otherwise : # e⁻ in ns orbital + # e⁻ in (n-1)d sublevel + # e⁻ in np sublevel

② Period #: corresponds to highest value of n.

③ Orbital-block : corresponds to the last orbital being occupied.

④ total # of electrons: sum of the number of electrons from each orbital (1s to last orbital occupied)

Question 8. Below is a valence orbital diagram of a fictional element Z that only has electrons in *s* and *p* orbitals. The valence orbital diagram is not correct.



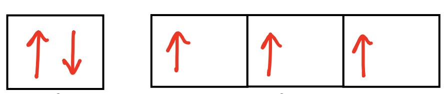
50*s* 50*p*

a) Identify the rule or principles (Hund's rule, Pauli's exclusion principle, and/or the aufbau principle) that were not followed when drawing Z's valence orbital diagram.

① 50s : • Pauli's exclusion principle not followed since both e⁻ in 50s have the same spin
• Pauli's exclusion principle requires that both e⁻ in an orbital have diff. spins.

② 50p : • Hund's rule not followed since not all orbitals in 50p are occupied by a single e⁻ before they are paired.
• Hund's rule says that each orbital should be singly occupied until it is no longer possible to do so

b) Draw the proper valence orbital diagram of element Z using the orbital boxes below.



50*s* 50*p*

c) Would the electron affinity of Z be positive or negative? Justify your answer.

Positive : ① High energy level (n = 50)

② The added electron will be added to the first 50p orbital box, disrupting an already stable configuration since 50p is half-filled.

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Level 4

← Fe^{2+} ← Fe^{3+}

Question 9. Iron (II) and iron (III) are two possible ions of iron. Which one do you believe is the most stable? Justify your answer.

Fe: 4s [↑↓] 3d [↑↓|↑|↑|↑|↑]

Fe^{2+}: 4s [ ] 3d [↑↓|↑|↑|↑|↑]

Fe^{3+}: 4s [ ] 3d [↑|↑|↑|↑|↑] → half-filled so electron configuration is stable, (more than Fe^{2+})

Question 10. A student is given an unknown element X that has a first ionization energy of 520 kJ/mol.

a) What does it mean for the element X to have an ionization energy of 520 kJ/mol?

- Require 520 kJ/mol of energy to remove a valence e^- from X in the gaseous state.

b) Write a thermochemical equation for element X.

$$X_{(g)} + 520\ kJ \rightarrow X^{+}_{(g)} + e^-$$

c) The student is given a second unknown element Y with a first ionization energy of 418.7 kJ/mol. Based on what you've learned in Chapter 3, give possible reasons why X and Y have different ionization energies.

① X has a lower n compared to Y

② X has a larger effective nuclear force.

③ Y 4s [↑] (removing e^- "closes" energy level n=4) X 4s [↑↓] → removing e^- will undo full sublevel.

④ Y 3d [↑↓|↑|↑|↑|↑] (removing e^- forms half-filled sublevel) X 3d [↑↓|↑↓|↑↓|↑↓|↑↓] → removing e^- will undo a full sublevel

d) Which element is more reactive? Justify your answer.

Based on the information available, Y would be more reactive since it requires less energy to remove a valence e^- (in the gaseous state).